Student Textbook page 480

Q1. The zinc atoms can reduce the copper(II) ions, even though the two metals are not in contact, because the electrons from the zinc travel through the electrodes via the connecting wire, before coming into contact with the copper(II) ions in solution.

Q2. The function of the salt bridge or porous cup in the voltaic cell is to allow movement of the excess anions and cations, preventing a build-up of charge in each half-cell. In the zinc-copper example, copper(II) ions are turned into copper atoms, leaving the sulfate ions behind. Over time, the negatively charged sulfate ions would build-up around the copper electrode. However, this is prevented by positively charged potassium ions in the salt bridge being attracted to the sulfate ions. Similarly, the chloride ions of the salt bridge are attracted to the increasing number of zinc ions.

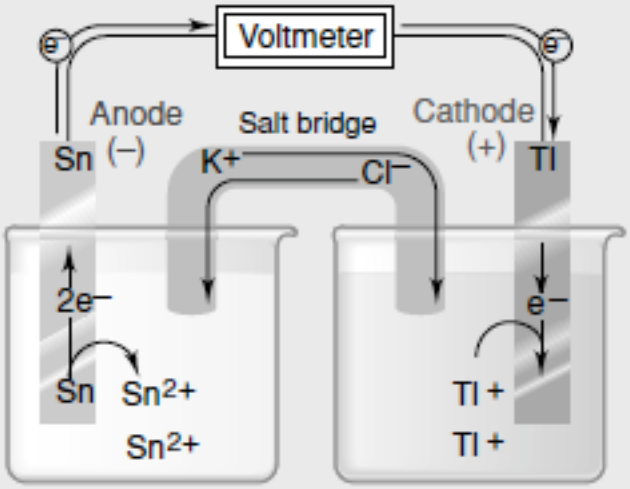
Q3. The main difference between the two half-cells is what happens at the electrodes. In the cathode, cations are converted into atoms, causing the electrode to gain in mass. Conversely, the electrode at the anode produces cations and consequently loses mass over time.

Student Textbook page 482

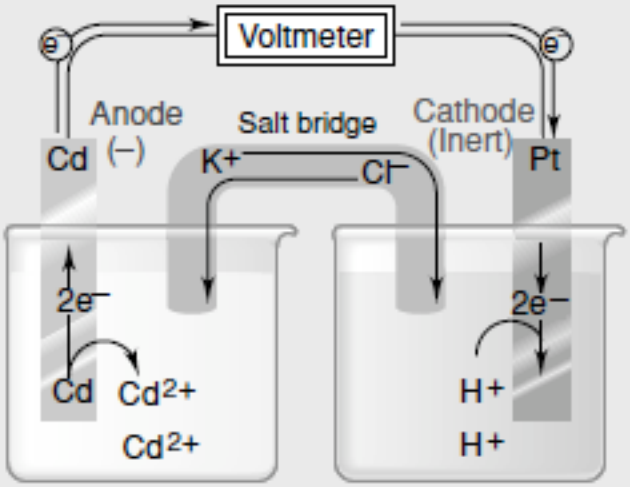
Q4. (a) In a test tube, the zinc would be the reducing agent and the copper(II) ions would be the oxidizing agent.

(b) In a Daniel cell, the copper(II) ions and the copper metal would be at the cathode while the zinc and zinc ions would be at the anode. Again students can determine this by remembering that electron flow is from anode to cathode, from reducing agent to oxidizing agent.

Q5. (a)



(b)



(a) anode: $Sn(s) \rightarrow Sn^{2+}(aq) + 2e^-$
cathode: $Tl^+(aq) + 1e^- \rightarrow Tl(s)$
Net: $Sn(s) + 2Tl^+(aq) \rightarrow Sn^{2+}(aq) + 2Tl(s)$

(b) anode: $Cd(s) \rightarrow Cd^{2+}(aq) + 2e^-$
cathode: $2H^+(aq) + 2e^- \rightarrow H_2(g)$
Net: $Cd(s) + 2H^+(aq) \rightarrow Cd^{2+}(aq) + H_2(g)$

Platinum is an inert electrode. A comma indicates that both species are in the same phase. There is a line instead of a comma to indicate that hydrogen gas is produced.

Q6. (a) The two half-reactions are:

$$MnO_4^-(aq) + 8H^+(aq) + 5e^- \rightarrow Mn^{2+}(aq) + 4H_2O(\ell)$$

$$2I^-(aq) \rightarrow I_2(s) + 2e^-$$

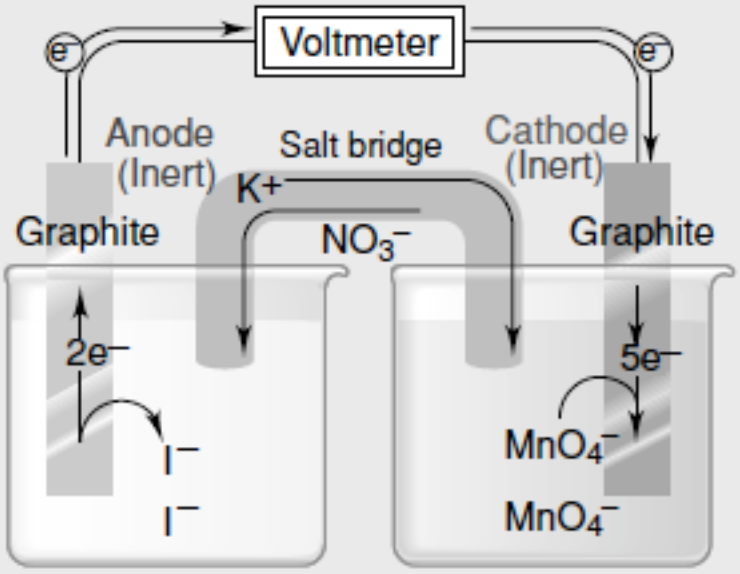
The overall balanced reaction is:

$$2MnO_4^-(aq) + 16H^+(aq) + 10I^-(aq) \rightarrow 5I_2(s) + 2Mn^{2+}(aq) + 8H_2O(\ell)$$

(b) The oxidizing agent is $MnO_4^-(aq)$ and the reducing agent is $I^-(aq)$.

(c) To form solid iodine from iodide ions, electrons must be lost (i.e. the ions must undergo oxidation). Since oxidation occurs at the anode, the solid iodine forms at the anode.

(d)



Q7. There is no double vertical line in the notation because there is no salt bridge or porous cup in the lemon.

Student Textbook page 487

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1. +0.29 V
2. +3.17 V
3. +0.14 V
4. +1.54 V

Student Textbook page 490

1. It is important to keep each half-cell separate so that an instantaneous reaction does not occur, "short-circuiting" the voltmeter.
2. The standard half-cell potentials are measured using the hydrogen half-cell as a reference. In the equation $E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode}$, the hydrogen half-cell is given a value of zero. Any measured voltage would therefore belong to the non-hydrogen half of the cell.
3. **(a)** $E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode}$
   $= -1.66\text{ V} - (-2.37\text{ V}) = +0.71\text{V}$

   (b) $E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode}$
   $= +2.87\text{ V} - (-2.93\text{ V}) = +5.80\text{ V}$

   (c) $E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode}$
   $= +1.23\text{ V} - (0.80\text{ V}) = +0.43\text{ V}$
4. **(a)** $E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode}$
   $= +0.34\text{ V} - (-0.26\text{ V}) = +0.60\text{V}$

   (b) $E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode}$
   $= +1.50\text{ V} - (-0.40\text{ V}) = +1.90\text{ V}$

   (c) $E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode}$
   $= -0.04\text{ V} - (-0.46\text{ V}) = +0.50\text{ V}$

5. No, a student using the Alberta Chemistry Data Booklet could not build a voltaic cell with a standard cell potential of 7.0 V. The strongest oxidizing agent $F_2(g)$ and the strongest reducing agent Li(s) would only yield 5.91 V. By building a gold-aluminium cell in series with two cells, a total of 7.32 V could be achieved.

6. The cell potential describes the potential difference between two electrodes of a cell, or the amount of energy on a charge as it moves between two electrodes. The cell potential is dependent on both the anode and the cathode used. The standard reduction potential is a measure of the amount of energy for only the reduction half of cell. Since reduction cannot happen without oxidation, reduction potentials are measured against a standard reference, the hydrogen half-cell, which is set at a reduction potential of 0.00V.

7. $E^{\circ}_{cell} = E^{\circ}_{cathode} - E^{\circ}_{anode}$
$= \text{“X” V} - (-0.76\text{ V}) = +1.75\text{ V}$
$X = 0.99\text{ V}$

Student Textbook page 493

Q8. The dry cell is similar to the Daniel cell in that both use a zinc anode. The dry cell differs from a Daniell cell in that the electrolyte has been thickened into a paste.

Q9. A battery consists of several cells connected in series.

Student Textbook page 496

Q10. The cell potential for the hydrogen fuel cell is:

$E^{\circ}_{cathode} - E^{\circ}_{anode} = E^{\circ}_{cell}$
$(1.23\text{ V}) - (0.00) = 1.23\text{ V}$

Q11. Typically a combustion reaction requires the presence of a flame, however, no flames are present in these reactions.

Q12. Currently, the source of hydrogen for the fuel cell comes from the conversion of ethane to ethene in the petrochemical industry. The cost of this process is one of the leading criticisms against the common use of hydrogen fuel cells.

Student Textbook page 500

Q13. (a) For a corrosion reaction:

$$E^{o}_{\text{cathode}} - E^{o}_{\text{anode}} = E^{o}_{\text{cell}}$$

$$(0.40\text{ V}) - (-0.45) = 0.85\text{ V}$$

(b) Students should be aware that the standard cell potentials are based on 1.0 mol/L concentrations and SATP conditions. If the concentrations or conditions of their rust "scenarios" differ, then they will get different (typically lower) cell potentials.

Q14. Aluminium provides cathodic protection on an iron object because it is a better reducing agent than iron. This means that the aluminium reacts first, saving the iron from rusting.

Q15. The corrosion problems most likely showed up in the Atlantic provinces due to high volumes of precipitation throughout the year (both rain and snow), coupled with the use of salt on winter roads. Taken together, these factors encourage corrosion.

Q16. (a) Given the cell potentials of the following half-reactions:

$$O_2(g) + 4H^+(aq) = E^{o}_{\text{cell}} = 1.23\text{ V}$$

$$O_2(g) + 2H^+(aq) = E^{o}_{\text{cell}} = 0.70\text{ V}$$

$$O_2(g) + 2H_2O(aq) = E^{o}_{\text{cell}} = 0.40\text{ V}$$

Elemental oxygen is a much better oxidizing agent under acidic conditions, providing an overall reaction of:

$2Fe(s) \rightarrow 2Fe^{2+}(aq) + 4e^-$

$O_2(g) + 4H^+(aq) + 4e^- \rightarrow 2H_2O(\ell)$

$E^o_{cathode} - E^o_{anode} = E^o_{cell}$

$(1.23\ V) - (-0.45) = 1.68\ V$

(b) Acid rain greatly increases oxygen's ability to act as an oxidizing agent and contributes to rusting.

Student Textbook page 501

1. The oxidizing agent is manganese dioxide. The reducing agent is zinc.

2. The top of a commercial 1.5-V dry cell is marked with a plus sign because the graphite cathode protrudes there.

3. (a) oxidation:

$Zn(s) + 2OH^-(aq) \rightarrow ZnO(s) + H_2O(\ell) + 2e^-$

reduction:

$Ag_2O(s) + H_2O(\ell) + 2e^- \rightarrow 2Ag(s) + 2OH^-(aq)$

overall cell reaction:

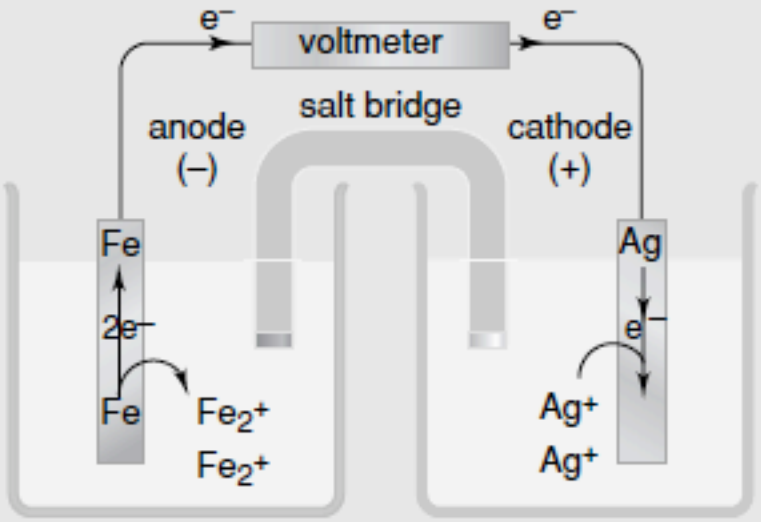
$Zn(s) + Ag_2O(s) \rightarrow ZnO(s) + Ag(s)$

(b) The anode is made of zinc, while the cathode is made of stainless steel.

4. In a flashlight, the two 1.5-V D batteries are connected in series and the voltages are additive; therefore, the total voltage is 3.0 V.

5. A 6-V dry cell battery contains four 1.5-V dry cells connected in series. When dry cells are connected in series, the voltages are additive, and $4 \times 1.5\ V = 6\ V$.

6. The potential of a cell depends only on the substances used for the anode and the cathode, not on the amounts of the substances used. A cell with a gold anode and an iron cathode would produce a cell potential of 1.95 V. If three of these cells were combined in series, a potential of almost 6 V would be produced. If only small quantities of these substances were used, the battery would be small. A large cell that produces only 1.5 V of potential contains larger quantities of the reactants required for the redox reaction. A large cell would be able to provide an electrical current over a longer period of time.
7. The zinc container of the dry cell is involved with the oxidation half-reaction. As zinc is oxidized, it forms ions and dissolves into the electrolyte paste. Therefore, the container has less and less zinc as the reaction progresses.
8. Assume that the cell represented by the shorthand notation is a typical galvanic cell.



Oxidation: $Fe(s) \rightarrow Fe^{2+}(aq) + 2e^-$

Reduction: $Ag^+(aq) + e^- \rightarrow Ag(s)$

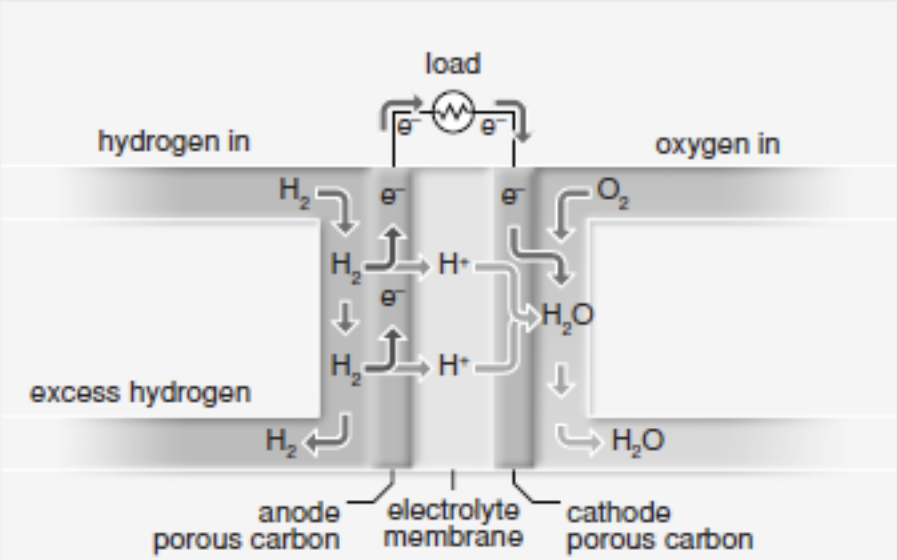
Overall cell reaction:

$Fe(s) + 2Ag^+(aq) \rightarrow Fe^{2+}(aq) + 2Ag(s)$

9. Both fuel cells and dry cells are alike in that they dispense with the need for an aqueous solution. In a dry cell, the electrolytes are in a paste, while fuel cells tend to rely on solid oxides, though the hydrogen fuel cell uses oxygen and hydrogen gases. One of the chief differences of fuel cells and dry cells, however, is that the main waste of fuel cells can be recycled, such as water in the case of the hydrogen cell, whereas the waste of many depleted batteries involve toxic heavy metal cations.

10. The main benefit of fuel cells for space use is the possibility of recycling waste to re-start, reducing the need to re-supply with fuel or the store and remove fuel waste. Second, many fuel cells operate at higher efficiencies, allowing for more energy to be produced overall. Finally, the production of water helps to reduce risks of fuel combustion.

11.



The anode reaction is: $H_2(g) \rightarrow 2H^+(aq) + 2e^-$

The cathode reaction is: $O_2(g) + 4H^+(aq) + 4e^- \rightarrow 2H_2O(\ell)$

The electricity is generated through the passing of electrons from the oxygen to the hydrogen in the presence of the platinum-coated carbon catalyst.

12. The use of salt lowers the melting point of snow, causing more water to be mobile and reactive. Furthermore, the salt also provides electrolytes required for the salt bridge.

13. When the steel cans corrode, they produce aqueous ions. However, the aluminum cans produce solid aluminium oxide, which slows down the corrosion process.

14. Zinc is a more reactive metal than iron, as noted by its cell potential, so it will react before iron does.

15. (a) Two metals that do not react in the presence of oxygen and water are silver and gold, as noted by their cell potentials being higher than that of the oxygen and water half-reaction.

(b) Since these metals are quite resistant to corrosion they are useful for industrial processes that take place under extremely adverse conditions that promote corrosion, such as high temperatures or acidic environments, like the aerospace or petroleum industry.

16. Silver sulfide forms when silver is exposed to small quantities of hydrogen sulfide found in the air.

$$4Ag(s) + 2H_2S(g) + O_2(g) \rightarrow 2Ag_2S(s) + 2H_2O(g)$$

The tarnish may be removed by an abrasive polish or, more easily, by chemically reducing the silver ions to silver atoms. To do this, place the tarnished piece of silver in a conducting solution, such as baking soda, in an aluminium pan. When the piece of silver comes in contact with the aluminium, the silver ions in the tarnish are reduced to silver.

$$3Ag_2S(s) + 2Al(s) \rightarrow 6Ag(s) + Al_2S_3(s)$$

The oxidation half-reaction is:

$Al(s) \rightarrow Al^{3+}(aq) + 3e^-$

The reduction half-reaction is:

$Ag^+(aq) + e^- \rightarrow Ag(s)$

Student Textbook page 508

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5. (a) +0.15 V

(b) Anode: Cd(s); Cathode: Ni(s)

(c) Oxidation: $Cd(s) \rightarrow Cd^{2+}(aq) + 2e^-$

Reduction: $Ni^{2+}(aq) + 2e^- \rightarrow Ni(s)$

Cell reaction: $Cd(s) + Ni^{2+}(aq) \rightarrow Ni(s) + Cd^{2+}(aq)$

6. (a) –0.15 V

(b) Anode: Ni(s) ; Cathode: Cd(s)

(c) Oxidation: $Ni(s) \rightarrow Ni^{2+}(aq) + 2e^-$

Reduction: $Cd^{2+}(aq) + 2e^- \rightarrow Cd(s)$

Cell reaction: $Ni(s) + Cd^{2+}(aq) \rightarrow Ni^{2+}(aq) + Cd(s)$

7. Co(s) and $Cl_2(g)$

8. Ag(s) and $O_2(g)$

Student Textbook page 510

Q17. (a) $2Cl^-(\ell) \rightarrow Cl_2(g) + 2e^-$

(b) $Ca^{2+}(\ell) + 2e^- \rightarrow Ca(s)$

(c) $Ca^{2+}(\ell) + 2Cl^-(\ell) \rightarrow Cl_2(g) + Ca(s)$

Q18. (a) $Li^+(el) + e^- \rightarrow Li(s)$

(b) $2Br^-(el) \rightarrow Br_2(\ell) + 2e^-$

(c) $2Li^+(el) + 2Br^-(el) \rightarrow Br_2(\ell) + 2Li(s)$

Q19. If alternating current were to supply the power to an electrolytic cell, then power would be supplied to both anode and cathode, causing the reverse reactions to occur as the power alternated.

Q20. Electrons leave the battery at the negative terminal and the cathode needs to accept electrons so that they can enter the solution containing the cations.

Student Textbook page 513

1. The four half-reactions involved are:

(1) $Br_2(\ell) + 2e^- \rightarrow 2Br^-(aq)$ $E^\circ = 1.07$ V

(2) $O_2(g) + 4H^+(aq) + 4e^- \rightarrow 2H_2O(\ell)$ $E^\circ = 0.81$ V (nonstandard reduction potential)

(3) $H_2O(\ell) + 2e^- \rightarrow H_2(g) + 2OH^-(aq)$ $E^\circ = -0.41$ V (nonstandard reduction potential)

(4) $Cu^+(aq) + e^- \rightarrow Cu(s)$ $E^\circ = 0.52$ V

The possible anode half-reactions are the reverse of (1) and (2). The possible cathode half-reactions are (3) and (4).

Using $E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode}$ the most positive (least negative) cell potential is obtained using half-reactions (2) and (4):

$E^\circ_{cell} = 0.52\text{ V} - 0.81\text{ V} = -0.29$ V

The overall reaction is:

$4Cu^+(aq) + 2H_2O(\ell) \rightarrow 4Cu(s) + O_2(g) + 4H^+(aq)$

The predicted products are copper, oxygen, and hydrogen ions.

2. (a) Since the electrolysis takes place in aqueous solution the relevant reduction potentials are:

(1) $Zn^{2+}(aq) + 2e^- \rightarrow Zn(s)$ $E^\circ = -0.76$ V

(2) $2H_2O(\ell) + 2e^- \rightarrow H_2(g) + 2OH^-(aq)$ $E = -0.41$ V

(3) $O_2(g) + 4H^+(aq) + 4e^- \rightarrow 2H_2O(\ell)$ $E = 0.81$ V

(4) $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$ $E^\circ = 0.34$ V

(1) and (2) are the possible half-reactions at the cathode. The reverse of (3) and (4) are the possible half-reactions at the anode. The lowest external voltage is required by the reaction including half-reaction (2) and the reverse of (4):

$$2H_2O(\ell) + Cu(s) \rightarrow Cu^{2+}(aq) + H_2(g) + 2OH^-(aq) \quad E_{cell} = -0.76 \text{ V}$$

The predicted products are hydrogen gas and copper (II) ions.

(b) Hydrogen gas requires overvoltage. Therefore, zinc deposits on the cathode.

3. (a) The half-reactions are:

Reduction (occurs at the cathode):

$Fe^{3+}(aq) + 3e^- \rightarrow Fe(s)$

Oxidation (occurs at the anode):

$2I^-(aq) \rightarrow I_2(s) + 2e^-$

The standard reduction potentials are:

$Fe^{3+}(aq) + 3e^- \rightarrow Fe(s)$ $E^\circ = -0.04$ V

$I_2(s) + 2e^- \rightarrow 2I^-(aq)$ $E^\circ = 0.54$ V

$$E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode}$$
$$= -0.04 \text{ V} - (0.54 \text{ V})$$
$$= -0.58 \text{ V}$$

The reaction has a negative standard cell potential; therefore, it is non-spontaneous under standard conditions.

(b) The half-reactions are:

Reduction (occurs at the cathode):

$Ag^+(aq) + e^- \rightarrow Ag(s)$

Oxidation (occurs at the anode):

$H_2SO_3(aq) + H_2O(\ell) \rightarrow SO_4^{2-}(aq) + 4H^+(aq) + 2e^-$

The standard reduction potentials are:

$SO_4^{2-}(aq) + 4H^+(aq) + 2e^- \rightarrow H_2SO_3(aq) + H_2O(\ell)$ $E^o = +0.17$ V

$Ag^+(aq) + e^- \rightarrow Ag(s)$ $E^o = +0.80$ V

$E^o_{cell} = E^o_{cathode} - E^o_{anode}$

$= 0.80\text{ V} - (0.17\text{ V}) = +0.63$ V

This reaction has a positive standard cell potential; therefore, it is spontaneous under standard conditions.

4. On discharging, the half-reactions are:

$Cd(s) + 2OH^-(aq) \rightarrow Cd(OH)_2(s) + 2e^-$

$NiO(OH)(s) + H_2O(\ell) + e^- \rightarrow Ni(OH)_2(s) + OH^-(aq)$

On recharging, therefore, the half-reactions are:

$Cd(OH)_2(s) + 2e^- \rightarrow Cd(s) + 2OH^-(aq)$

$Ni(OH)_2(s) + OH^-(aq) \rightarrow NiO(OH)(s) + H_2O(\ell) + e^-$

To balance charge, multiply the oxidation equation by 2:

$2Ni(OH)_2(s) + 2OH^-(aq) \rightarrow 2NiO(OH)(s) + 2H_2O(\ell) + 2e^-$

The overall reaction for recharging a nicad battery can then be written as:

$Cd(OH)_2(s) + 2Ni(OH)_2(s) \rightarrow Cd(s) + 2NiO(OH)(s) + 2H_2O(\ell)$

5. According to the caption for Figure 13.27 on page 511, a lead-acid car battery consists of six 2-V cells. The total potential is 12 V. Therefore, an external voltage of over 12 V is required to recharge the battery.

6. The reaction in an electrolytic cell is non-spontaneous; it requires an input of energy to occur. The energy is in the form of electrical energy. The external source of electrons forces electrons onto one electrode, which becomes negative relative to the other electrode. This drives the movement of ions and non-spontaneous redox reactions.

7. **(a)** The question is asking whether the following reaction is spontaneous:

$$Al(s) + 3H_2O(\ell) \rightarrow H_2(g) + Al(OH)_3(s)$$

The oxidation of aluminium is considered to occur at the anode. The reduction of water is considered to occur at the cathode.

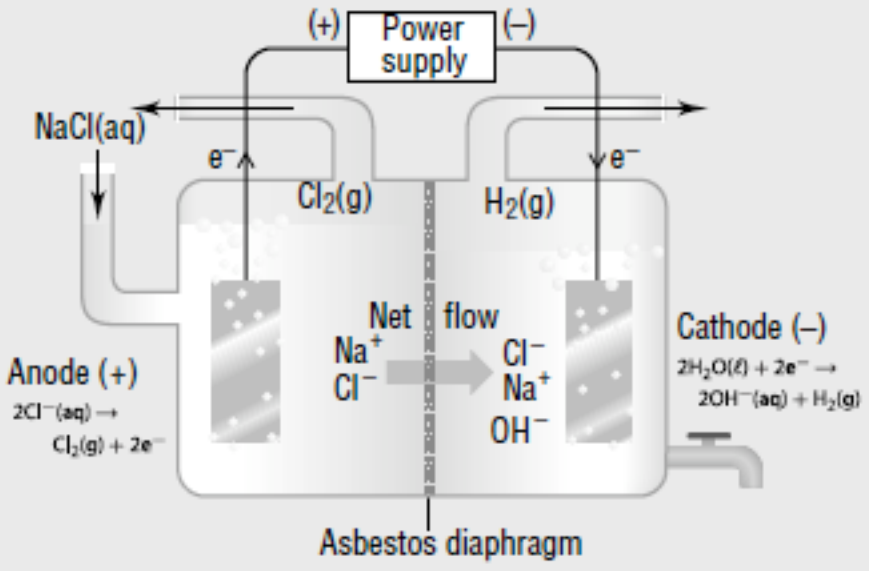
$$E^o_{cell} = E^o_{cathode} - E^o_{anode}$$
$$= -0.83\ V - (-1.66\ V) = +0.83\ V$$

The reaction should be spontaneous under standard conditions.

(b) The boiling point of water is 100 °C, which is not the standard temperature. The aluminium vessel is likely covered by a thin but durable coating of unreactive aluminium oxide. Also, predictions are not reliable when reduction potentials are similar (aluminium is just below water on the chart of standard reduction potentials), and when gaseous products are expected, which is the case in this example.

8. There are advantages and disadvantages to lead-acid batteries. Lead-acid batteries have been available for almost 150 years, so the technology is stable and has been improved on over the years. Modern batteries are completely sealed to minimize one of the potential risks, which is the possibility of a chemical burn if someone comes in contact with the acid in the battery. Although the charging process is never perfect, lead-acid batteries are long-lasting and if handled properly, much of the lead (up to 97%) can be recycled during disposal. On the negative side, lead is heavy, which decreases the energy efficiency of the car, and if disposed of improperly, lead will contaminate soil and water. If lead-acid batteries are cracked, the sulfuric acid inside will leak and may cause burns. In addition, if improperly charged or if an electrical short-circuit takes place, hydrogen gas that is produced when the battery is charged could be ignited if sparked, leading to a fire or an explosion.

9.



The chlor-alkali cell is an electrolytic cell that oxidizes chloride ions, $Cl^-(aq)$ from brine (salt water) and reduces water, forming hydrogen gas and hydroxide ions.

Oxidation: $2Cl^-(aq) \rightarrow Cl_2(g) + 2e^-$
Reduction: $2H_2O(l) + 2e^- \rightarrow H_2(g) + 2OH^-(aq)$

The salt water enters the cell, and a battery provides the potential required to electrolyze the salt water. The products of the electrolysis are removed from the cell – the chlorine gas from the anode, the hydrogen gas from the cathode, and the aqueous sodium hydroxide is isolated from the electrolyte.

Student Textbook page 516

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9. 2.97 g

10. 8.90 min

11. 1.85 A

12. (a) 1.09 g increase

(b) Because the overall balanced equation can be determined from the given information, the mole ratio is sufficient information to calculate the mass change.

Student Textbook page 520

1. (a) Nickel is electroplated onto the silver master disc. The master disc must provide electrons to reduce the nickel ions to nickel atoms. Reduction reactions take place at the cathode of the electrolytic cell. Therefore, the master disc is the cathode.

(b) Number of moles of Ni formed

$$= \frac{1 \text{ g Ni}}{58.69 \text{ g/mol}}$$

$$= 1.704 \times 10^{-2} \text{ mol}$$

To reduce each mole of nickel, 2 mol of electrons are required. Therefore, 0.034 08 mol of electrons are required for 1.704×10^{-2} mol of nickel.

$$\text{Quantity of electricity} = 0.034\,08 \text{ mol } e^- \times \frac{96\,500 \text{ C}}{1 \text{ mol } e^-}$$
$$= 3.289 \times 10^3 \text{ C}$$

2. $$\text{Quantity of electricity} = 3.00 \times 10^4 \text{ A} \times 8.64 \times 10^4 \text{ s}$$
$$= 2.59 \times 10^9 \text{ C}$$
$$\text{Amount of electrons} = 2.59 \times 10^9 \text{ C} \times \frac{1 \text{ mol } e^-}{96\,500 \text{ C}}$$
$$= 2.69 \times 10^4 \text{ mol } e^-$$

The half-reactions are:

$Na^+(\ell) + 1e^- \rightarrow Na(s)$

$2Cl^-(\ell) \rightarrow Cl_2(g) + 2e^-$

Amount of sodium formed

$$= 2.69 \times 10^4 \text{ mol } e^- \times \frac{1 \text{ mol Na}}{1 \text{ mol } e^-}$$
$$= 2.69 \times 10^4 \text{ mol Na}$$

Mass of sodium formed

$$= 2.69 \times 10^4 \text{ mol Na} \times \frac{22.99 \text{ g}}{\text{mol Na}}$$
$$= 6.18 \times 10^5 \text{ g}$$
$$= 618 \text{ kg}$$

Amount of chlorine formed

$$= 2.69 \times 10^4 \text{ mol } e^- \times \frac{1 \text{mol } Cl_2}{2 \text{ mol } e^-}$$
$$= 1.34 \times 10^4 \text{ mol } Cl_2$$

Mass of chlorine formed

$$= 1.34 \times 10^4 \text{ mol } Cl_2 \times \frac{70.90 \text{ g}}{\text{mol } Cl_2}$$
$$= 9.50 \times 10^5 \text{ g}$$
$$= 950 \text{ kg}$$

3. Time in seconds

$$= 365 \frac{\text{days}}{\text{year}} \times 24 \frac{\text{hours}}{\text{day}} \times \frac{60 \text{ min}}{\text{hour}} \times \frac{60 \text{ s}}{\text{min}}$$
$$= 3.154 \times 10^7 \text{ s}$$

Quantity of electricity

$= 2.00 \times 10^2\ \text{A} \times 3.154 \times 10^7\ \text{s}$

$= 6.31 \times 10^9\ \text{C}$

Amount of electrons

$$= 6.31 \times 10^9\ \text{C} \times \frac{1\ \text{mol}\ e^-}{96\,500\ \text{C}}$$

$= 6.54 \times 10^4\ \text{mol}\ e^-$

Assume the reduction reaction is:

$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$

Amount of copper formed

$$= 6.54 \times 10^4\ \text{mol}\ e^- \times \frac{1\ \text{mol Cu}}{2\ \text{mol}\ e^-}$$

$= 3.27 \times 10^4\ \text{mol Cu}$

Mass of copper formed

$$= 3.27 \times 10^4\ \text{mol Cu} \times \frac{63.6\ \text{g}}{\text{mol Cu}} = 2.08 \times 10^6\ \text{g}$$

$= 208\ \text{t}$

4. Even with the cost of transporting the bauxite to Canada, it is still cheaper to extract the aluminium by electrolytic processes because of the abundance of relatively cheap electricity in Canada.

5. $$m\text{Al(s)} = 22.7\ \text{t} \times \frac{1000\ \text{kg}}{1\text{t}} \times \frac{1000\ \text{g}}{1\ \text{kg}}$$

$$= 2.27 \times 10^7\ \text{g Al(s)}$$

$$n\text{Al(s)} = \frac{m}{M}$$

$$= \frac{2.27 \times 10^7\ \text{g Al(s)}}{\frac{26.98\ \text{g Al(s)}}{\text{mol Al(s)}}}$$

$8.41 \times 10^5\ \text{mol Al(s)}$

$Al(s) \rightarrow Al^{3+}(aq) + 3e^-$

$$n_e = 8.41 \times 10^5\ \text{mol Al(s)} \times \frac{3\ \text{mol}\ e^-}{1\ \text{mol Ag(s)}}$$

$$= 2.52 \times 10^6\ \text{mol}\ e^-$$

$$\text{amount of electrons} = 2.52 \times 10^6 \text{ mol } e^- \times \frac{9.65 \times 10^4 \text{ C}}{\text{mol } e^-}$$
$$= 2.43 \times 10^{11} \text{ C}$$

$$q = I\Delta t$$
$$\Delta t = \frac{2.43 \times 10^{11} \text{ C}}{2.5 \times 10^5 \text{ C/s}}$$
$$= 9.74 \times 10^5 \text{ s}$$
$$= 11 \text{ days}$$

It would take 11 days of continuous cell operation to produce 22.7 t of aluminium.

$$E = qV$$
$$= (2.44 \times 10^{11} \text{ C})(4.5 \text{ V})$$
$$= 1.1 \times 10^{12} \text{ J}$$

The amount of energy used to produce 22.7 t of aluminium is 1.1×10^{12} J.

6. Nickel extraction is analogous to copper extraction. The $Cu_2S(s)$ is roasted in air and some of it is converted to $Cu_2O(s)$:

$$2Cu_2S(s) + 3O_2(g) \rightarrow 2Cu_2O(s) + 2SO_2(g)$$

The copper compounds react to produce copper:

$$Cu_2S(s) + 2Cu_2O(s) \rightarrow 6Cu(s) + SO_2(g)$$

The copper is further refined by electrolysis. Sulfur dioxide gas reacts with water to form sulfurous acid, which contributes to acid precipitation. To prevent the release of sulfur dioxide into the atmosphere, various recovery programs have been initiated. One example is the electrostatic precipitation program at INCO's refinery in Sudbury, Ontario.

Student Textbook pages 522–523

Answers to Understanding Concepts Questions

1. (a) Electrodes are the conductors that carry electrons into and out of the cell. Oxidation and reduction occur at the electrodes.

(b) Electrolytes are substances that dissolve in water to form ions that can move and conduct electricity.

(c) The external voltage is a source of electricity that is included in the external circuit of an electrolytic cell.

2. $Mg(s)\,|Mg^{2+}(aq)||Cd^{2+}(aq)|\,Cd(s)$

3. The oxidation half-reaction is:

$$NO(g) + 2H_2O(\ell) \rightarrow NO_3^-(aq) + 4H^+(aq) + 3e^-$$

The reduction half-reaction is:

$$I_2(s) + 2e^- \rightarrow 2I^-(aq)$$

Balance the electrons and add the equations to get the overall cell reaction:

$$2NO(g) + 4H_2O(\ell) + 3I_2(s) \rightarrow 2NO_3^-(aq) + 8H^+(aq) + 6I^-(aq)$$

4. The hydrogen electrode is defined as the reference electrode against which other reduction potentials are measured. The reduction potential of the standard hydrogen electrode is assigned a value of 0.00 V.

5. Electrolyzing aqueous solutions often results in unwanted products due to redox reactions involving water.

6. The two half-reactions are:

$$CuI(s) + 1e^- \rightarrow Cu(s) + I^-(aq)$$

$$N_2O(g) + 2H^+(aq) + 2e^- \rightarrow N_2(g) + H_2O(\ell)$$

First assume that copper is oxidized at the anode.

Write a balanced equation.

$$2Cu(s) + I^-(aq) \rightarrow 2CuI(s) + 2e^-$$

Add the two equations to get the balanced net ionic equation.

$$N_2O(g) + 2H^+(aq) + 2Cu(s) + 2I^-(aq) \rightarrow N_2(g) + H_2O(\ell) + 2CuI(s)$$

$E^o_{cell} = E^o_{cathode} - E^o_{anode} = 1.77\ V - (-0.18\ V)$

$= +1.95\ V$ (spontaneous reaction)

Now assume that copper ions are reduced at the cathode:

$$N_2(g) + H_2O(\ell) + 2CuI(s) \rightarrow N_2O(g) + 2H^+(aq) + 2Cu(s) + 2I^-(aq)$$

$E^o_{cell} = E^o_{cathode} - E^o_{anode} = -0.18\ V - (+1.77\ V)$

$= -1.95\ V$ (non-spontaneous reaction)

7. (a) The oxidizing agent is $PbO_2(s)$ and the reducing agent is Pb(s).

(b) The oxidizing agent is Pb(s) and the reducing agent is $PbO_2(s)$.

8. In order of the most effective to least effective oxidizing agents, based on the table of standard reduction potentials and assuming that the reduction potential for cobalt is between that of hydrogen ions and zinc (based on the activity series), the order is:

$Br_2(\ell)$, $H^+(aq)$, $Co^{3+}(aq)$, $Zn^{2+}(aq)$

9. Based on the table of standard reduction potentials, the order is:

Al(s), $H_2(g)$, Ag(s), $Cl^-(aq)$

10. The metals would plate-out in order of their standard reduction potentials. Of the three metals, silver has the most positive standard reduction potential (i.e., it is most easily reduced). It will plate out first. Then copper will plate out, followed by iron.

Answers to Applying Concepts Questions

11. (a) $Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$ (anode)

$Fe^{2+}(aq) + 2e^- \rightarrow Fe(s)$ (cathode)

$E^o_{cell} = E^o_{cathode} - E^o_{anode} = -0.45\ V - (-0.76\ V)$

$= +0.31\ V$ (spontaneous)

(b) $Cr(s) \rightarrow Cr^{3+}(aq) + 3e^-$ (anode)

$Al^{3+}(aq) + 3e^- \rightarrow Al(s)$ (cathode)

$E^o_{cell} = E^o_{cathode} - E^o_{anode} = -1.66\ V - (-0.74\ V)$

$= -0.92\ V$ (non-spontaneous)

(c) $2H_2O_2(\ell) \rightarrow O_2(g) + 2H^+(aq) + 2e^-$ (anode)

$Ag^+(aq) + e^- \rightarrow Ag(s)$ (cathode)

$E^o_{cell} = E^o_{cathode} - E^o_{anode} = 0.80\ V - (+0.70\ V)$

$= +0.10\ V$ (spontaneous)

12. (a) Connect a voltmeter to the external circuit with the negative terminal connected to the tin anode and the positive terminal connected to the lead cathode.

(b) This cell is unlikely to find any practical uses because the cell voltage is so low and the components are heavy. ($E^o_{cell} = +0.01\ V$)

13. (a) oxidation at anode: $Fe(s) \rightarrow Fe^{2+}(aq) + 2e^-$

reduction at cathode: $Ag^+(aq) + e^- \rightarrow Ag(s)$

$E^o_{cell} = E^o_{cathode} - E^o_{anode}$

$= 0.80\ V - (-0.45\ V) = +1.25\ V$

Balance charge and add the two half-reactions to get the cell reaction:

$2Ag^+(aq) + Fe(s) \rightarrow 2Ag(s) + Fe^{2+}(aq)$

(b) oxidation at anode: $Ag(s) \rightarrow Ag^+(aq) + e^-$

reduction at cathode: $Fe^{2+}(aq) + 2e^- \rightarrow Fe(s)$

$E^o_{cell} = E^o_{cathode} - E^o_{anode} = 0.80\ V - (-0.45\ V)$

$= +1.25\ V$ (spontaneous)

Balance charge and add the two half-reactions to get the cell reaction:

$2Ag(s) + Fe^{2+}(aq) \rightarrow 2Ag^+(aq) + Fe(s)$

(c) For the voltaic cell, the mass of the iron electrode will decrease as iron atoms form iron(II) ions. The mass of the silver electrode will increase as the silver ions form silver atoms.

(d) For the electrolytic cell, the mass of the silver electrode will decrease as silver atoms form silver ions. The mass of the iron electrode will increase as iron(II) ions become iron atoms.

14. (a) Students should suggest using inert electrodes to electrolyze the aqueous zinc bromide and testing the solutions around the electrodes (e.g., with phenolphthalein) to determine what is produced. Bromine in solution would appear orange-yellow.

(b) Based on standard reduction potentials, you would not expect zinc to be oxidized. Instead, you would expect water to be oxidized.

15. Numerous combinations are possible. For example, a cell in which $Al^{3+}(aq)$ is reduced to Al(s) and Fe(s) is oxidized to $Fe^{3+}(aq)$ has a cell potential of 1.62 V. Five of these cells in series make a battery of 8.10 V.

16. (a) Luigi Galvani's great interest was "animal electricity," which he studied in his post as a teacher of medicine in Bologna. He constructed a crude electric cell with two different metals and the natural fluids from a dissected frog. Alessandro Volta's first important invention was the electrophorus, a device capable of passing an electric charge to another object. Arguably his most important contribution was the development and study of the "voltaic pile," which provided proof that electricity could be conducted through, and be produced by, non-organic matter.

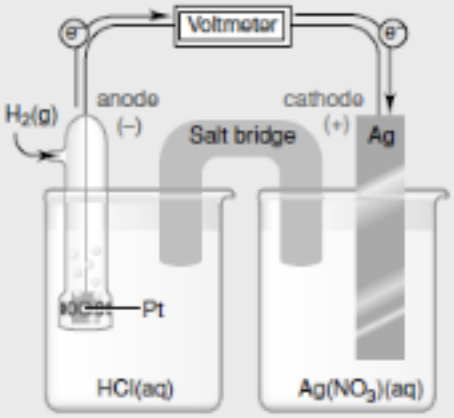
(b) Davy studied Galvani's work with the "voltaic pile," and followed the examples of Nicholson and Carlisle by using electrolysis to extract elemental potassium and sodium from molten potash and table salt, respectively. He continued his work on electrolysis and he discovered that any electrolyte would pass current, as long as it was capable of oxidizing one of the metals. He also determined that the intensity of the effect (the voltage) was directly related to the reactivity of the electrolyte with the metal.

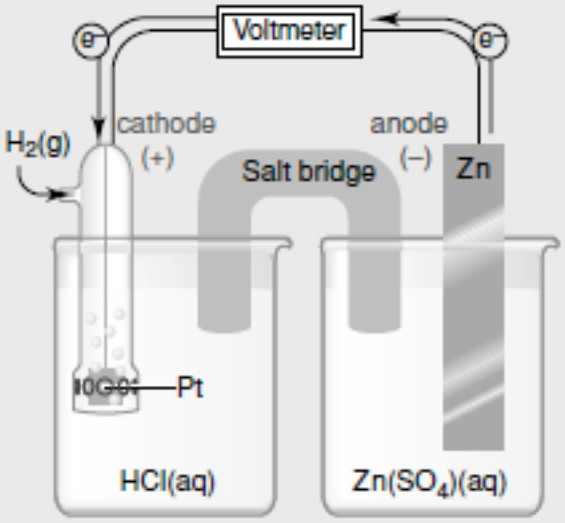
In the 1830s Faraday worked on developing his ideas on electricity. He stated a new theory of electrochemical action. One of the results of this work was that he coined, with William Whewell, many of the words we currently use, including electrode, electrolyte, anode, cathode, and ion. His work led him to reject the traditional theory that electricity was an imponderable fluid or fluids, and he proposed that electricity was a form of force that passed from particle to particle of matter.

17. The surface of the moon does not have an atmosphere or water. Since oxygen is not available for the redox reaction, iron would not corrode.

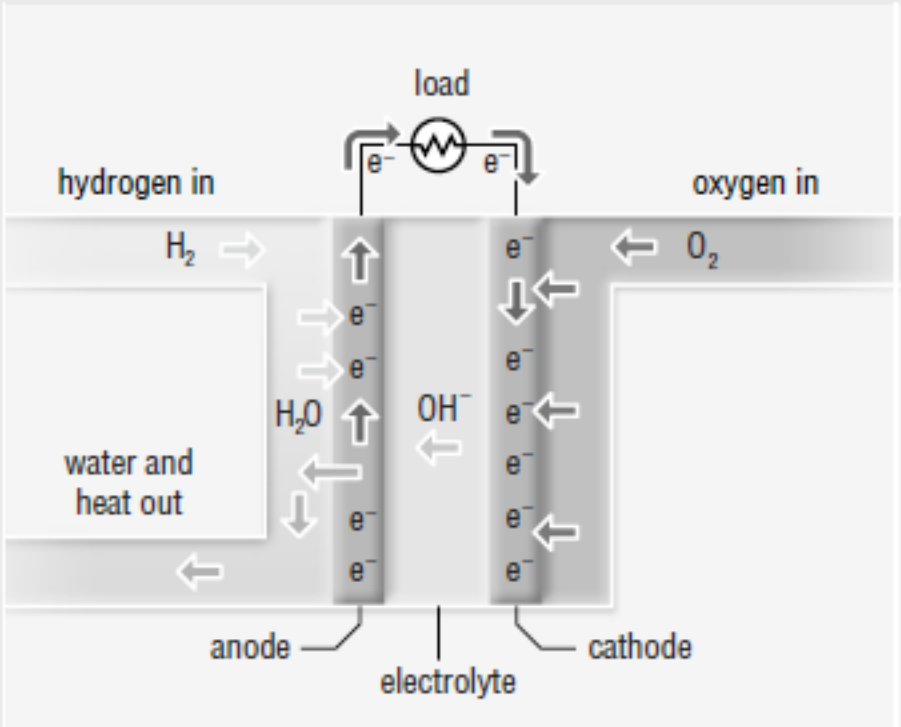
18. It is simply reversing the direction of the electron flow so that a reduction reaction becomes an oxidation reaction, and visa versa.

19.




Voltmeter
e⁻
e⁻
cathode
(+)
anode
(–)
Salt bridge
Zn
H2(g)
Pt
HCl(aq)
Zn(SO4)(aq)

20.


load
e⁻
e⁻
hydrogen in
oxygen in
H2
O2
e⁻
e⁻
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e⁻
e⁻
H2O
OH⁻
water and
heat out
anode
electrolyte
cathode

Answers to Solving Problems Questions

21. 2.46 V; Al/Al^{3+} oxidation half cell; Ag/Ag^{+} reduction half cell

22. 1.2 g of copper

23. Time in s $= 55.0 \text{ min} \times 60 \text{ s/min} = 3.300 \times 10^3 \text{ s}$

$$\text{Quantity of electricity} = 3.65 \text{ A} \times 3.300 \times 10^3 \text{ s}$$
$$= 1.204 \times 10^4 \text{ C}$$

$$\text{Amount of electrons} = 1.2048 \times 10^4 \text{ C} \times \frac{1 \text{ mol } e^-}{96\,500 \text{ C}}$$
$$= 0.125 \text{ mol } e^-$$

Assume the reduction reaction is:

$Mg^{2+}(aq) + 2e^- \rightarrow Mg(s)$

$$\text{Amount of Mg formed} = 0.1248 \text{ mol } e^- \times \frac{1 \text{ mol Mg}}{2 \text{ mol } e^-}$$
$$= 0.0624 \text{ mol Mg}$$

$$\text{Mass of Mg formed} = 0.0624 \text{ mol Mg} \times \frac{24.31 \text{ g}}{1 \text{ mol Mg}}$$
$$= 1.52 \text{ g}$$

24. $$\text{Number of moles Na} = 1000 \text{ g} \times \frac{1 \text{ mol}}{22.99 \text{ g}}$$
$$= 43.50 \text{ mol}$$

$$\text{Number of moles Al} = 1000 \text{ g} \times \frac{1 \text{ mol}}{26.98 \text{ g}}$$
$$= 37.06 \text{ mol}$$

The amount of each material in terms of moles is very close. However, reducing 1 mole $Al^{3+}(aq)$ requires 3 mole of electrons, while reducing 1 mole $Na^{+}(aq)$ requires only 1 mole of electrons. Producing 1 kg of aluminium would require about 2.6 times more electricity than producing 1 kg of Na(s).

25. 1.98 kg/h

26. 0.133 A

Answers to Making Connections Questions

27. The quantity of electricity depends on the current generated. The current generated depends on the quantity and identity of substances involved in the electrochemical reaction. Because the batteries are different sizes, they probably contain different materials. Also, the larger battery is likely to last longer. Even if the batteries produced the same current, the larger battery would produce a greater overall quantity of electricity.

28. (a) Yes, because aluminium is reduced more readily than iron. The aluminium would be oxidized, forming a relatively inert oxide coating, and thus hold the iron gutter in place while protecting the iron from oxidation.

(b) No, because after the aluminium siding formes the surface oxide, the iron nails would rust and flake to pieces.

29. A cell fuelled by aluminium generates power through an electrochemical reaction between the metal, placed in a saline or alkaline solution, and oxygen from the air. Electricity is produced as the aluminium oxidizes (at the anode), so it acts as the reducing agent. Oxygen from the air acts as a cathode. To recharge the cell when the aluminium is consumed, the plates are replaced and more electrolyte is added.

30. About 95% less energy is required to process used aluminium than is to process the aluminium from the ore. The aluminium in most products tends to be relatively pure, and there is little surface oxide to remove during the recycling process.

31. Students may say they are against the smelter because the waste products that are formed, such as sulfur dioxide, can contribute to acid precipitation. Students may say they are in favour of the plant because of the economic situation and the fact that the modern mechanisms for removing waste (sulfur dioxide) from the roasting process results in a reduction in the quantities of these pollutants. Accept all reasoned and reasonable answers, but students should demonstrate some understanding of the processes involved.

32. The iron "ribs" of the statue, which supported the copper "skin," corroded due to the effects of wind, salty air, pollution, and time. Because the iron was in contact with the copper, which acted as a cathode, the ribs rusted much faster than if they had not been in contact with the copper. During the restoration process, the iron ribs were removed and replaced with copper ribs.

33. **(a)** Estimates will vary based on your classroom. Ensure students consider all uses of batteries, including watches, cellular phones, and portable CD players.

(b) Students may suggest purchasing rechargeable or longer-life batteries as a partial solution. They could suggest using an AC adapter for battery powered devices whenever possible. Or, students could simply find ways to reduce their use of battery-powered devices.